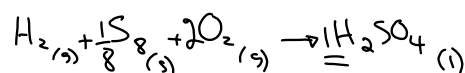
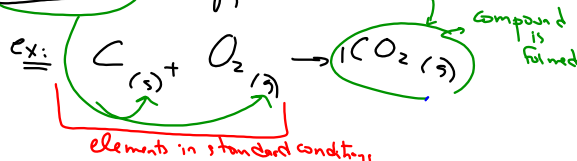


Hess's Law part 2:

Standard energy of formation:



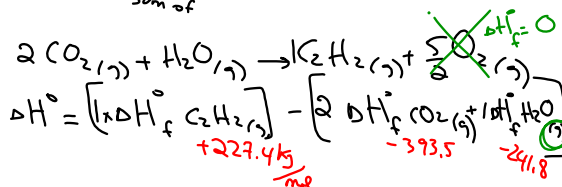
Given as: $\Delta H_f^\circ \text{CO}_{2(g)} = -393.5 \text{ kJ/mol}$

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Reactants \rightarrow Products $\Delta H_f^\circ = ?$

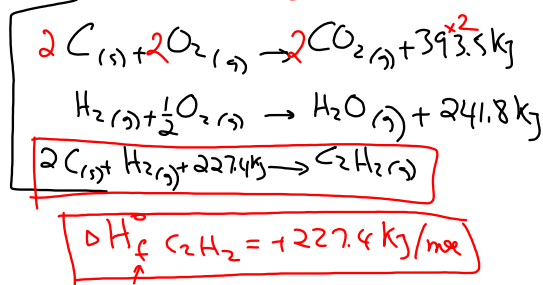
$$\Delta H^\circ = \sum \Delta H_f^\circ \text{ products} - \sum \Delta H_f^\circ \text{ reactants}$$

↑
sum of



$$= 227.4 - [2(-393.5) + (-241.8)]$$

$$= 1256.2 \text{ kJ}$$



Things to watch out for:

ΔH_f° : formation of compound from the elements in standard state

ΔH_f° elements in standard form = 0

ex: $\Delta H_f^\circ \text{I}_2(s) = 0$

$\Delta H_f^\circ \text{I}_2(l) \neq 0$

Handout
P254

Elements in standard form:

$\text{Na}_{(s)}, \text{Br}_{2(l)}, \text{P}_{4(s)}, \text{S}_{8(s)}$

Pay attention to the physical states of substances

use mole coefficients when determining ΔH_f°
need a balanced equation.

Question 1:

Combustion of methane:

 $\Delta H^\circ_{\text{Combustion CH}_4} \Rightarrow \text{Using } \Delta H^\circ_f$

$$\Delta H^\circ_{\text{comb}} = [1 \times \Delta H^\circ_f(\text{CO}_2(g)) + 2 \Delta H^\circ_f(\text{H}_2\text{O}(g))] - [1 \times \Delta H^\circ_f(\text{CH}_4(g)) + 2 \times \Delta H^\circ_f(\text{O}_2(g))]$$

$$= [2(-241.8) + (-393.5)] - (-74.6)$$

$$= -802.5 \text{ kJ/mol}$$

$$Q_{\text{water}} = mc\Delta T$$

$$= (300\,000) \text{ g} \times 4.18 \frac{\text{J}}{\text{g}^\circ\text{C}} \times 52.5^\circ\text{C}$$

$$= 65\,835\,000 \text{ J}$$

$$= 65\,835 \text{ kJ}$$

$$Q_{\text{sys}} = -Q_{\text{sur}}$$

$$\frac{1 \text{ mol CH}_4}{-802.5 \text{ kJ}} = \frac{x \text{ mol CH}_4}{-65\,835}$$

$$\text{or } \frac{Q}{n} = \Delta H \quad n = \frac{\Delta H}{Q}$$

$$\text{or } n = \frac{Q}{\Delta H} \Rightarrow \frac{-65\,835 \text{ kJ}}{-802.5 \text{ kJ/mol}}$$

$$n = 82.04 \text{ mol}$$

$$m = n \times M \quad \text{or} \quad \frac{82.04 \text{ mol}}{x \text{ g}} = \frac{1 \text{ mol}}{16.04 \text{ g}}$$

$$= 82.04 \times 16.04 \text{ g}$$

$$= 1316.7 \text{ g}$$

$$= 1.32 \text{ kg}$$